Name ____ AP Chemistry

Electrochemistry Lab

Introduction

Chemical reactions are commonly classified as acid-base, precipitation, complexation or redox reactions. Since redox reactions comprise the largest number of known reactions, it is important to know how to recognize and understand these chemical transformation processes. Redox reactions are reactions that involve a transfer of electrons. Oxidation refers to the loss of electrons by a molecule, atom, or ion. Reduction refers to the gain of electrons by a molecule, atom or ion.

Common examples of redox reactions include corrosion of iron or steel and the patina that forms on bronze statues. Redox reactions always involve a reactant that undergoes oxidation and a reactant that undergoes reduction. The reactant undergoing oxidation is called a reducing agent whereas the reactant undergoing reduction is called an oxidizing agent. Redox reactions are essential for life and play a central role in respiration, photosynthesis, and many other important metabolic pathways.

Redox reactions are classified by whether they occur spontaneously. Spontaneous redox reactions are called galvanic reactions whereas non-spontaneous redox reactions are called electrolytic reactions. Gibb's free energy (ΔG) must be negative for any spontaneous reaction and negative for any non-spontaneous reaction. The potential developed for any cell is related thermodynamically to free energy as:

$$\Delta G° = -nFE°_{cell}$$

where n is the number of moles of electrons being transferred, F is 96,485 C/mole of electrons (Faraday's constant), and $E°_{cell}$ is the potential developed between the two half reactions at standard conditions. Thus a positive $E°_{cell}$ is required for a spontaneous reaction while a negative $E°_{cell}$ is required for a non-spontaneous reaction.

Many times, redox reactions are conducted under nonstandard conditions. Under these conditions, E_{cell} is not the same as $E°_{cell}$. The Nernst equation enables the prediction of the cell potential under any situation so long as the concentrations of reactants and products are known:

$$E_{cell} = E°_{cell} - \frac{RT}{nF} \ln \frac{[products]^x}{[reactants]^y}$$

where R is the thermodynamic gas constant (8.314 J/Kmol), T is temperature in kelvin, n is the number of moles of electrons transferred in a balanced reaction, and F is Faraday's constant. Solids are not included in the natural log term; concentrations of the products and reactants are raised by their stoichiometric coefficients.

The activity series is a series of metals, in order of reactivity from highest to lowest. It is used to determine the products of single replacement reactions, whereby metal A will replace another metal B in a solution if A is higher in the series. Here are two single replacement reactions:

$$Mn(s) + AuCl_2(aq) \rightarrow MnCl_2(aq) + Au(s)$$

$$Au(s) + MnCl_2(aq) \rightarrow Mn(s) + AuCl_2(aq)$$

They are mirror images of each other but which one will actually occur? If we look on an activity series chart we would find that Mn is listed as more reactive (higher on the chart) than Au. That means that Mn metal will replace any metal, such as Au that is below it on the activity chart, from a compound such as $AuCl_2$. Thus Mn will end up in the compound $MnCl_2$ and force Au to precipitate out as Au solid. In the two reactions above, the top reaction will take place and not the bottom one.

First in this lab, you will be measure the volt of several galvanic cells and then use the Nernst equation to calculate the cell potential in the nonstandard conditions. Then during the virtual lab you will place a solid metal (Au for example) and a second metal nitrate ($Mn(NO_3)_2$ as an example) together and see if they react. If they do, then the solid metal used in the combination should be placed higher on your activity series list then the metal in the nitrate. If no reaction occurs then the solid metal should be placed lower on the activity series list.

Prelab Questions

1) A redox reaction occurs between copper metal and Ag^+ ions.
   (a) Write a balanced redox equation for this reaction.
   (b) Calculate the standard potential ($E°_{cell}$) for this reaction from the standard half-reaction table.
   (c) Is this reaction spontaneous? Explain.
   (d) Calculate the nonstandard potential (E_{cell}) for this reaction when $[Cu^{2+}]$ = 1 M and $[Ag^+]$ = 0.1 M.
2) Predict the activity, from weakest to strongest, of the following metals: Cu, Fe, Pb, Zn, Ag, and Al.

Procedure

Part A: Determining Cell Potentials

1) Obtain a well plate and add 2 mL (two squirts of a disposable pipette) of the following solutions in the appropriate cells:

A1	0.1 M $Zn(NO_3)_2$	B1	0.1 M $Cu(NO_3)_2$
A2	0.1 M $Zn(NO_3)_2$	B2	0.1 M $AgNO_3$

A3	0.1 M $Zn(NO_3)_2$	B3	0.1 M $Fe(NO_3)_3$
A4	0.1 M $Zn(NO_3)_2$	B4	0.1 M $Mg(NO_3)_3$
A5	0.1 M $Zn(NO_3)_2$	B5	0.1 M $Pb(NO_3)_3$
C1	0.1 M $Cu(NO_3)_2$	D1	0.1 M $Mg(NO_3)_3$
C2	0.1 M $Fe(NO_3)_3$	D2	0.1 M $Cu(NO_3)_2$
C3	0.1 M $Fe(NO_3)_3$	D3	0.1 M $AgNO_3$
C4	0.1 M $Mg(NO_3)_3$	D4	0.1 M $Pb(NO_3)_3$
C5	0.1 M $Pb(NO_3)_3$	D5	0.1 M $Cu(NO_3)_2$
C6	0.1 M $Cu(NO_3)_2$	D6	0.1 M $AgNO_3$

2) Obtain copper foil, iron nail, lead foil, magnesium foil, silver foil, and zinc strip.
3) Place a piece of paper towel in a solution of KNO_3. Once it is soaked, place it between cells A1 and B1 so it is connecting the two solutions.
4) Using the voltmeter, connect one lead to a piece of zinc and the other to a piece of copper. Place each metal in the appropriate solution.
5) Measure and record the voltage in Data Table 1. Note: the voltage should be positive so you may need to switch the leads if the voltage is negative.
6) Repeat step 3-5 for the following pairs of cells: A2-B2, A3-B3, A4-B4, A5-B5, C1-D1, C2-D2, C3-D3, C4-D4, C5-D5, and C6-D6. Use a new strip of paper towel (soaked in KNO_3 solution) for each pair and use the appropriate piece of metal for each pair.
7) Record all voltage in Data Table 1.
8) Rinse out well plate in the waste beaker.

Part B: Activity Series

1) Go to www.eduweblabs.com. Choose "Advanced Chem" at the top of the page and then click on "Activity Series" in the series of labs.
2) Use the log-in information provided and follow the procedure in the virtual lab.
3) Record all data in Data Table 2.

Data Tables

Data Table 1

Cell #	Metal/Solution	Cell #	Metal/Solution	Measured Voltage (V)
A1		B1		
A2		B2		
A3		B3		
A4		B4		
A5		B5		
C1		D1		
C2		D2		
C3		D3		

C4		D4		
C5		D5		
C6		D6		

<u>Data Table 2</u>

		Solutions					
Metals							

Calculations

1) For each of the reactions below, give the anode and cathode half-reaction:

Cell Pairing	Reaction Equation	Measured Voltage (V)	n	Calculated Voltage (V)	Percent Error
A1-B1	Anode: Cathode:				
A2-B2	Anode: Cathode:				
A3-B3	Anode: Cathode:				
A4-B4	Anode: Cathode:				
A5-B5	Anode: Cathode:				
C1-D1	Anode: Cathode:				
C2-D2	Anode: Cathode:				
C3-D3	Anode: Cathode:				
C4-D4	Anode: Cathode:				
C5-D5	Anode:				

	Cathode:				
C6-C6	Anode: Cathode:				

2) Based on the half-reactions, determine the balanced number of electrons transferred. Write these under "n" in the table above.
3) Using the Nernst equation, calculate the cell potential for each reaction.
4) Calculate the percent error for each reaction using the following formula:

$$\% \, Error = \frac{Measured\ Voltage - Calculated\ Voltage}{Calculated\ Voltage} x\ 100$$

Postlab Questions

1) Describe the difference between the measured voltages and calculated voltages using the percent error in your explanation. Which reactions had the largest percent error? Smallest percent error?
2) Why did you or would you get a negative voltage for a galvanic cell?
3) Rank the metals from the activity series from most active to least active. Use observations as evidence to defend your ranking.